Worksheet: Balancing, mole:mole ratios, mass-mass, mass-mole, mass-molecule problems

1. For the following:

a. Balance the equation

b. State the mole:mole ratios

c. State the mass in grams of each substance

a. P4 + O2  P4O6

b. Fe(OH)3 + HCl  FeCl3 + H2O

2. Phosphorus trichloride (PCl3) is made when white phosphorus (P4) reacts with chlorine gas: P4 + Cl2  PCl3

a. Balance the equation

b. How many grams of phosphorus are required to produce 5.49 g PCl3?(1.24 g)

3. How many moles of oxygen are required to prepare 142 g P4O10 from elemental white phosphorus?

a. Balance the equation

b. Calculate moles of oxygen. (2.5 moles)

c. What mass of oxygen is needed for this reaction to occur? (80 g)

4. Suppose 2.17 g HgO is thermally decomposed to elemental mercury and oxygen.

a. Balance the equation

b. What mass of mercury will be produced? (2.01 g)

c. How many oxygen molecules will be produced? (3.02 x 1021 molecules)

5. How many grams of Cu2S could be produced from 9.90 g CuCl reacting with H2S gas? (Hint: This is a double replacement reaction).

a. Balance the equation

b. Find grams of Cu2S (7.96 g)

Worksheet: Mass-Mass Problems, Theoretical Yield(#1,2,4,5,8)

1. How many grams of hydrogen can be produced from the reaction of 11.5 g of sodium with an excess of water? (0.505 g H2)

2. An excess of nitrogen reacts with 2.0 g of hydrogen. How many grams of ammonia are produced? (11.2 g NH3)

3. How many grams of oxygen are required to burn completely 85.6 grams of carbon?

(228 g O2)

4. In problem 3, how many grams of CO2 will be formed? (314 g CO2)

5. In the decomposition of potassium chlorate, 64.2 grams of oxygen are formed. How many grams of potassium chloride are produced? (99.7 g KCl)

6. The action of carbon monoxide on iron(III) oxide can be represented by the equation:

Fe2O3 + 3CO  2Fe + 3CO2. What would be the minimum amount of carbon monoxide used if 18.7 g of iron were produced? (14.1 g CO)

7. How many grams of hydrochloric acid are required to react completely with 75.1 grams of calcium hydroxide? (74.0 g HCl)

8. How many grams of hydrogen are produced when 5.62 grams of aluminum reacts with excess hydrochloric acid? (0.631 g H2).

Worksheet: mass-volume, and volume-volume

1. What volume of hydrogen at STP can be produced from the reaction of 6.54 grams of zinc with hydrochloric acid? (single replacement rxn) (2.24 L)

2. How many grams of sodium chloride can be produced by the reaction of 112 ml of chlorine at STP with excess sodium? (composition rxn) (0.585 g)

3. An excess of hydrogen reacts with 14 grams of nitrogen. How many liters of ammonia will be produced at STP? (composition rxn) (22.4 L)

4. How many liters of oxygen are required to burn 1.00 liter of methane, CH4? (combustion rxn) (2 L)

5. How many liters of carbon dioxide will be produced by burning completely 5.00 liters of ethane, C2H6? (combustion rxn) (10 L)

6. What volume of oxygen is required to burn completely 401 ml of butane, C4H10? (combustion rxn) (2.61 L)

Worksheet: Mass-mass, mass-volume or volume-mass, and volume-volume

Also, 1)classify the following reactions(m-m, m-v, v-v and 2) determine which of the following are theoretical yield problems?

1. How many grams of NaCl will be produced when 22.85 g of HCl are neutralized by an excess of NaOH? (36.63 g)

2. What volume of hydrogen gas is produced when 135 grams of aluminum are completely reacted with excess sulfuric acid(hydrogen sulfate) at STP? (168.23L)

3. How many grams of magnesium oxide are produced when 10.0 grams of magnesium burn in an excess of oxygen? (16.6 g)

4. How many liters of ammonia gas are produced when 35 grams of liquid nitrogen(in this experiment) completely react with excess hydrogen at STP? (56.0L)

5. How many grams of aluminum would react completely with 17.5 grams of copper (II) chloride? (2.34 g)

6. What volume of bromine gas(in this experiment) is produced if 75.2 L of chlorine react with excess hydrogen bromide at STP? (75.2 L)

Percent and Theoretical Yield

1. A student was preparing copper metal by the reaction of 1.274 g of copper(II) sulfate with zinc metal. She isolated a yield of 0.392 g of copper. What was her theoretical and percent yield of copper? (0.5072 g Cu, 77.3% yield)

2. A student reacted 3.22 grams of sodium bicarbonate with an excess of hydrochloric acid. They produced 2.35 grams of sodium chloride. The gases carbon dioxide and water vapor were also produced. What was the student’s theoretical yield and percent yield of sodium chloride? ( 2.24 g, 104.9% yield)

3. A lab group reacted 0.75 grams of nails(iron) with copper(II) chloride to produce copper and iron(II) chloride. The lab group produced 0.77 grams of copper. Calculate the theoretical yield and percent yield of copper. (0.85 g,

90.6% yield)

4. A lab group reacted 0.92 grams of nails(iron) with copper(II) sulfate to produce copper and iron(II) chloride. This lab group produced 1.01 grams of copper. Calculate the theoretical yield and percent yield of copper. ( 1.05 g Cu; 96.2% yield)

\*\*\*5. A sample of lime, CaO weighing 69 g was prepared by heating 131 g of limestone(95% CaCO3). Carbon dioxide is also produced. What was the theoretical yield and percent yield of the reaction? (69.72 g CaO; 99% yield)

Worksheet: Limiting Reactants

1. If 6.57g iron reacts with 10.7 g HCl, then H2 and iron(II) chloride are produced. Determine which reactant is in excess? How much should have been put in? How much was wasted(in excess)? Determine the mass of each product. Show that the law of conservation of matter applies in this reaction!! (HCl is in excess. 8.57 g HCl should have been put in. 2.13 g HCl was wasted(in excess). 0.2376 g H2 , 14.91 g FeCl2)

2. When 8.76 g Al reacts with 29.37 g HCl, then aluminum chloride and hydrogen are produced. Which reactant is in excess? How much should have been put in? How much was wasted(in excess)? Calculate the mass of each product. Show that the law of conservation of matter applies to this reaction. (Al is in excess. 7.24 g should have been put in. 1.52 g of Al was wasted(in excess). 35.8 g AlCl3 and 0.8136 g H2 )

Worksheet: Limiting Reactants

1. 24 g of ammonia(NH3) reacts with 35 g of hydrogen chloride in a composition reaction forming ammonium chloride. How much ammonium chloride is formed? If any reactant remains unreacted, how much is left over? How much should have been put in? Show that the law of conservation of matter applies in this reaction. (NH3 is in excess. 16.36 g should have been put in. 7.64 g was in excess. 51.36 g of ammonium chloride is produced)

2. 13 g of CH4 is burned in 50 g of oxygen. How many grams of CO2 and H2O are produced? Which reactant is left over and by how much(Note: it is not in excess by very much)? ( CH4 is in excess. 12.54 g should have been put in. 0.46 g were in excess. 34.38 g CO2 and 28.16 g H2O are produced)

3. 196 g of sulfuric acid(H2SO4) reacts with 316 g of calcium acetate in a double replacement reaction. How many grams of each product are produced? If any reactant remain unreacted, which one is it and how much is left? (Note: sometimes there is not an excess as is the case in this problem. So…. You can use either reactant to do the mass-mass problems). ( 240.04 g HC2H3O2 and 272.06 g CaSO4 )

Harder Mass-Mass Problems

1. How many grams of air are required to complete the combustion of 93 g of phosphorus to diphosphorus pentaoxide assuming the air to be 23% oxygen by mass. (522 g air)

2. How many grams of carbon dioxide can be produced from the combustion of 1000kg of coke that is 90% carbon? (3,300,000g CO2)

3. What mass of a sample that is 98% sulfur would be required in the production of 75 kg of H2SO4 by the following reaction sequence: (~24,500 g unpure sulfur)

S8 + 8O2  8SO2

2SO2 + O2  2SO3

SO3 + H2O  H2SO4

\*\*\*\*4. How many pounds of 58% pure salt cake (Na2SO4) could be produced from150 lbs of 85% pure salt. (267.15 lbs unpure salt cake)

2NaCl + H2SO4  Na2SO4 + 2HCl

Harder Mass-Mass Problems

1. Calculate the number of grams of lime(CaO) that can be prepared by heating 200 grams of limestone that is 95% pure CaCO3. Carbon dioxide is also produced. (106.4 g CaO)

2. Commercial sodium hydrosulfite is 90% pure Na2S2O4. How many grams of the commercial product could be made by using 100 grams of zinc? The reactions follow: (295.9 grams of commercial product, unpure)

Zn + 2SO2  ZnS2O4

ZnS2O4 + Na2CO3  ZnCO3 + Na2S2O4.

3. Bi + 4HNO3 + 3H2O  Bi(NO3)3 .5H2O + NO

a. How many grams of Bi(NO3)3.5H2O would be formed from a solution of 10.4 grams of bismuth in nitric acid? (24.14 g bismuth nitrate pentahydrate)

b. How many grams of 30% nitric acid(containing 30% HNO3 by mass) is required to react with 10.4 g of bismuth? (41.82 g unpure nitric acid)

\*\*\*4. 2NaCl + H2SO4  Na2SO4 + 2HCl

How many pounds of 83.4% pure salt cake (Na2SO4) could be produced from 250 lbs. of 94.5% pure salt? (344.25 lbs unpure sodium sulfate)

Helpful conversions:

453.6 g=1 lb

2000 lbs=1 ton

Worksheet: Everything!!!

1. A compound was known to be either CuCl2 or CuBr2. A 5.00 g sample yielded 2.36 g of copper. What was the compound?

2. Calculate the formula of a compound, given that 55.85 g of iron combines with 32.06 g of sulfur.

3. How much iron could be obtained from 1 ton of iron ore containing 45% Fe2O3? (Put your answer in pounds of Fe)

4. How much SO2 could be obtained from burning 25 g sulfur in oxygen? (Put your answer in moles of SO2)

5. What would be the mass of the residue if 8.375 g of U(SO4)2.9H2O were heated until the water had evaporated?

6. The element M forms the chloride MCl2. This chloride contains 75% chlorine. Calculate the atomic mass of M.

7. How many tons of Fe2O3 will contain 12 tons of Fe?

8. A compound is either ZnBr2 or ZnI2. An 8.00 g sample yielded 1.64 g of zinc. What is the compound?

9. How many pounds of KCl will be formed if 50 pounds of KClO3 are decomposed by heating?

10. It was found that 10.0 g of a pure compound contains 3.65 g K, 3.33 g of Cl, and 3.02 g of O. Calculate the empirical formula of the compound. Its molecular mass is 106.55. What is its molecular formula?

Answers:

1. CuCl2 6. 23.65 g

2. FeS 7. 17.16 tons

3. 629.29 lbs. 8. ZnI2

4. 0.78 moles 9. 30.4 lbs.

5. 6.082 g 10. KClO2, KClO2

Title: Mole Relationship in a Chemical Reaction Lab

Purpose: In this experiment, you will test the Law of Conservation of Matter by causing a reaction to occur with a given amount of reactant. You will then determine the mass of one of the products. You will then compare the experimental and theoretical mol:mol ratio between one of the reactants and one of the products. Additionally, you will determine which reactant is in excess and which is limiting. Once the excess reactant is determined, you will determine how much should have been used and how much was wasted(in excess). Then, you will compare the experimental yield of sodium chloride compared to the theoretical yield of sodium chloride produced. Finally, you will determine the percent yield of sodium chloride. This will be used to determine your accuracy grade.

Directions:

1. Clean an evaporating dish and rinse it with water from the sink. Dry it thoroughly with paper towels.
2. Obtain the mass of the evaporating dish and the watch glass to the nearest 0.01 grams.
3. With a scoopula, add about 3 grams of sodium bicarbonate (NaHCO3) to the evaporating dish and read the mass to the nearest 0.01 grams.
4. Obtain about 8 mL of 6M hydrochloric acid in a clean graduated cylinder. Record the exact measurement.
5. Slowly add the acid to the sodium bicarbonate from your graduated cylinder. Allow the drops to enter the lip of the evaporating dish so that they flow down the side gradually.
6. Continue adding all of the acid slowly. Make sure to keep the watch glass on.
7. Tilt the dish from side to side to make sure the liquid has reached the entire solid.
8. Heat the liquid in the evaporating dish over a Bunsen burner. Make sure to keep the watch glass on top of the evaporating dish. Be careful of splatter. You should be using a blue flame and waving it underneath the evaporating dish back and forth.
9. Make sure it appears to be excessively dry. Heating should take approximately 10 to 15 minutes.
10. Remove the heat from under the dish and let it cool.
11. Record the new mass to the nearest 0.01 grams.

Data:

1. Mass of empty evaporating dish + watch glass \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

2. Mass of dish + watch glass + NaHCO3 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

3. Mass of NaHCO3 \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

4. Mass of dish + watch glass + NaCl \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

5. Mass of NaCl \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

6. Volume of HCl \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Calculations/Analysis Questions:

1. Write the balanced equation that occurred in this experiment.

2. What is the theoretical mol:mol ratio between the sodium bicarbonate and the sodium chloride.

3. Calculate the experimental mass of sodium bicarbonate used and sodium chloride produced.

4. Change the mass to moles to determine the experimental mol:mol ratio. Compare this to the theoretical. How close did you come? Did you prove the Law of Conservation of Matter? Was matter created or destroyed?

5. Now take the volume of HCl and using the idea of molarity to calculate the number of moles of HCl used. Notice the theoretical mol:mol ratio between the sodium bicarbonate and HCl Compare this to the experimental mol:mol ratio between the sodium bicarbonate and HCl and determine which was in excess. Was HCl in excess? Now calculate how much HCl should have been used comparing it to the moles of sodium bicarbonate that was used. How much HCl was wasted(in excess) in ml ?

6. Now using the limiting reactant(sodium bicarbonate), calculate the theoretical yield of sodium chloride that should have been produced.

7. Using the experimental and theoretical values of sodium chloride, calculate the percent yield of sodium chloride produced.

For the “A”

8. Research sodium bicarbonate, hydrochloric acid, and sodium chloride. Find physical properties, chemical properties, and uses for each of these compounds.

Chemistry - Nail Lab

# Iron(II) or (III) – Copper (II) chloride Reaction. Iron (II) or (III) – Copper (II) sulfate reaction.

## Purpose

The purpose is to determine the ratio of copper produced to iron consumed in a single replacement reaction. Determine the type of iron(?)chloride formed in this reaction. Is it iron +2 or iron +3?

## Procedure

**Day 1**

1. Label, then mass a 150 mL beaker.

2. Put between 6.0 and 8.0 g of copper(II) chloride crystals in the beaker OR put between 3.50 and 4.00 g of copper (II) sulfate in the beaker. Check with your teacher as to which chemical we are using this year. But, you must record the exact amount you are using.

3. Add about 50 mL distilled water to the beaker. Stir to dissolve the solid with a stirring rod.

4. Mass 2 or 3 nails together to ± 0.01g.

5. Place the nails in the copper(II) chloride solution. Your teacher will show you how to do this so the reaction will continue to occur overnight. Observe the reaction; record your observations. Place the labeled beaker in the place designated by your teacher.

**Day 2**

6. Remove the nails. Rinse and scrape all the precipitate (copper metal) from the nails into your labeled 150 mL beaker. Clean the nails with steel wool and mass them again to determine the mass lost by the nails. Return the nails back to the teacher.

7. Decant solution from the 150 mL beaker. Rinse the precipitate with about 25 mL of distilled water. Let the copper settle to the bottom and then decant trying to lose as little of the solid copper as you can. Your teacher will show you how to do this. Rinse with distilled water three times making sure to decant each time but keeping as much of the copper as possible. Make sure you rinse and decant at least three times. Then place the labeled beaker in the designated area to dry overnight.

8. What are we trying to rinse away so we can get just the copper product??

### Day 3

9. Mass the beaker + dry copper. Discard the copper in the place designated by your teacher. Wash your beaker and let dry.

## Data:

|  |  |
| --- | --- |
| Mass 150 mL beaker |  |
| Mass 150 mL beaker + copper(II) chloride |  |
| Mass nails before reaction |  |
| Mass nails after reaction |  |
| Mass 150 mL beaker + dry copper |  |

## Calculations:

1. Determine the mass of copper produced and the mass of iron used during the experimental reaction.

2. Calculate the moles of iron and moles of copper involved in the reaction.

3. Determine the experimental ratio moles of iron : moles of copper

4. State the two possible reactions that could occur by writing two balanced equations. Iron reacting with the copper(II)chloride(or copper(II)sulfate). One represents iron +2 and the other represents iron +3. Determine which reaction occurred by comparing the experimental mol:mol ratio above to the theoretical mol:mol ratio from the balanced equation.

5. Determine the amount of copper(II) chloride crystals(or copper(II)sulfate) that should have reacted with the amount of iron that actually reacted. Perform a mass-mass problem starting with mass of iron changing to mass of copper(II) chloride crystals(or copper(II) sulfate depending on which chemical used). Make sure you use the correct theoretical mol:mol ratio for this conversion.

6. Determine how much copper(II) chloride(copper(II)sulfate) crystals were wasted by subtraction. Initial copper crystals used – copper crystals that actually reacted = copper crystals wasted.

7. Determine the theoretical amount of copper that should have been made(Theoretical yield). Again, perform another mass-mass problem starting with grams of iron and calculating the grams of copper.

8. Determine what was the percent yield of copper. (Experimental yield of copper / theoretical yield of copper) x 100

**Conclusion:**

1. Why did the reaction stop? Which reactant was used up? How do you know?

2. Describe what was happening to the atoms of iron and copper during the reaction. What is this type of reaction called?

3. What would happen to the ratio of iron to copper if you had placed more nails in the beaker?

What would happen to the ratio of iron to copper if you let the reaction go for less time?

4. What is the accepted ratio of iron atoms to copper atoms in this reaction.? Account for differences between your experimental value and the accepted value.

For the “A”

Research iron, copper, copper(II) chloride (or copper(II) sulfate if that was used). State physical properties, chemical properties and uses for each of the substances.

PRACTICE TEST-STOICHIOMETRY

1. How many grams of carbon are needed to burn completely to produce 80 liters of carbon dioxide at STP?

Balanced equation: (2 pts)

Show work: (6 pts)

Classify:\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (2 pts)

2. What volume of methane, CH4, is necessary to burn in order to produce 50 liters of water at STP? (combustion reaction)

Balanced equation: (2 pts)

Show work: (6 pts)

Classify:\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (2 pts)

3. How many grams of magnesium oxide will be produced if 9.5 grams of magnesium reacts with oxygen?

Balanced equation (2 pts)

Show work: (6 pts)

Classify:\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (2 pts)

4. A compound was known to be either CuCl2 or CuBr2. A 5.00 g sample yielded 2.36 g of copper. What was the compound?

5. When 40 grams of magnesium burns in 10 grams of oxygen, then magnesium oxide is produced.

a) Write the balanced equation (2 pts)

b) Which reactant is limiting and which is excess? Show all work! (6 pts)

c) How much of the excess should have been put in? (2 pts)

d) How much excess was wasted? (2 pts)

e) How much magnesium oxide is produced? Note: this is the theoretical yield of magnesium oxide and you need to do a mass-mass problem. (6 pts)

f) If a student did this experiment and produced 15 grams of magnesium oxide, what would be the percent yield for this experiment? (2 pts)

g) State the law of conservation of matter and energy (2 pts)

h) Show that the law of conservation of matter is proved in this reaction mathematically. (2 pts)

6. What is the molarity of a MgCl2 solution that is made from 450 grams of magnesium chloride in 5.5 liters of solution? (6 pts)

7. Suppose a propane tank, C3H8 , for grilling was filled with 380 grams of propane. How much air will react with the propane if the air has 22 % oxygen.

(Hint: combustion reaction: propane + oxygen yields carbon dioxide + water)

(10 pts)

8. Note the following sequence of reactions: (10 pts)

N2 + 3H2 → 2NH3

NH3 + HCl → NH4Cl

If you have 950 grams of air that is 77% nitrogen, how much ammonium chloride can be produced based on the above sequence of reactions?